Titration

Acid + base ⟶ salt

a titrant: Precised Known Concentration

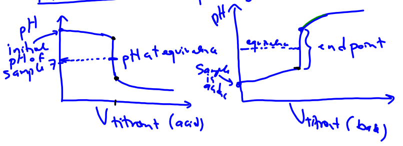
a sample: unknown concentration

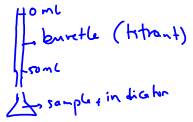
an indicator: changes color with pH

equivalence point: time when the amount of the titrant (in moles) added is stoichiometrically equivalent to the amount of the sample (in moles)
"Acid reacted completely with the base"!

end point: when indicator changes color.

titration curve: pH against volume of titrant





3 possible equivalence pH:

< 7	7	> 7
usually (5-7)	strong acid + strong base	usually (7-9)
strong acid + weak base		strong base + weak acid

Recall: $n_a C_a V_a = n_b C_b V_b$

Qualitatively (predict) pH at equivalence:

$HNO_3 + NH_3 \rightarrow$

acid: strong acid
base: weak base } acidic pH < 7

Quantitatively: P543

111) pH at equivalence when 50.00 mL of 0.150 mol/L ammonia solution is titrated with 0.150 mol/L HCl(aq).

titration equation: $NH_{3(aq)} + HCl_{(aq)} \leftrightarrow NH_4^+{}_{(aq)} + Cl^-{}_{(aq)}$
apply Bronsted-Lowry definitions

Organize all the concentrations in a table:

	NH_3 +	HCl ⇌	NH_4^+ +	Cl^-
(mol) n	C×V	C×V		
(L) V	0.0500	$C_1V_1=C_2V_2$		
(mol/L) C	0.150	0.150		

(if ratio same as n_{acid} = n_{base}) $n = C \times V$ $C = \frac{n}{V}$ $C_1V_1 = C_2V_2$

	NH_3 +	HCl ⇌	NH_4^+ +	Cl^-
(mol) n	0.007500	0.007500	0.007500	0.007500
(L) V	(0.05000) +	(0.05000) =	0.1000 (total volume)	0.1000
(mol/L) C	0.150	0.150		
	50.00		0.07500	0.07500

Salt: $NH_4^+ + Cl^-$
only NH_4^+ will affect pH. (it came from a weak base)
Cl^- doesn't affect (it came from strong acid)

Hydrolysis of NH_4^+:

	NH_4^+ +	H_2O ⇌	NH_3 +	H_3O^+
I	0.07500		0	0
C	−x		+x	+x
E	0.07500−x		x	x

K_a since H_3O^+ is produced:

$K_a\ NH_4^+$ (remember $K_aK_b = K_w$)
$K_b\ NH_3 = 1.8 \times 10^{-5}$

$K_a\ NH_4^+ = 10^{-14} \div (1.8 \times 10^{-5})$
$= 5.6 \times 10^{-10}$

use approximation method: $\frac{0.075\ mol/L}{K_a} > 1000$ ✓

x^2 ... 5.6×10^{-10} ... $x = 6.48 \times 10^{-6}$ mol/L